Weak Acids, Weak Bases and Buffers

1) Write the Brønsted-Lowry reaction for weak acid HCN reacting with H_2O.

   $HCN + H_2O \rightleftharpoons H_3O^+ + CN^-$

2) Write the Brønsted-Lowry reaction for weak base NH_3 reacting with H_2O

   $NH_3 + H_2O \rightleftharpoons OH^- + NH_4^+$

3) Using the information from the Chemical Equilibria Exercises, write the equilibria expression for the acid HCN in 1) above using the substitution $K_a = K$

$$K_a = \frac{[H_3O^+][CN^-]}{[HCN]}$$

4) Write the equilibrium expression for the base NH_3 in 2) above using the substitution $K_b = K$.

$$K_b = \frac{[OH^-][NH_4^+]}{[NH_3]}$$

5) Calculate the pH of a solution which is 0.010 M in HCN.
The K_a of HCN is 4.2×10^{-10}.
$HCN + H_2O \rightleftharpoons H_3O^+ + CN^-$
— x—>

	[HCN]	$[H_3O^+]$	$[CN^-]$
initially	0.010	0	0
equilibrium	0.010 - x	x	x

substituting: $4.2\times10^{-10} = \frac{x^2}{0.010 - x}$ and assume $x << 0.010$

$x^2 = 4.2 \times 10^{-12}$ so: $x = 2.0 \times 10^{-6}$

Ans: 5.69

6) Calculate the pH of a solution which is 0.10 M in HCN and 0.10 M in NaCN. The K_a of HCN is 4.2×10^{-10}. (Notice that the important chemical species here is the CN^- ion. Na^+ is a spectator ion)
$HCN + H_2O \rightleftharpoons H_3O^+ + CN^-$
— x—>

	[HCN]	$[H_3O^+]$	$[CN^-]$
initially	0.010	0	0
equilibrium	0.10 - x	x	0.10 + x

substituting: $4.2\times10^{-10} = \frac{(0.10 - x)x}{0.10 - x}$ and assume $x << 0.10$

$x = 4.2 \times 10^{-10}$

Ans: 9.38

7) Calculate the pH of a solution which is 0.10 M in HCN and 0.30 M in NaCN.
The K_a of HCN is 4.2×10^{-10}. (What is Na^+? Ans: spectator ion)
$HCN + H_2O \rightleftharpoons H_3O^+ + CN^-$
— x—>

	[HCN]	$[H_3O^+]$	$[CN^-]$
initially	0.010	0	0
equilibrium	0.10 - x	x	0.30 + x

substituting: $4.2\times10^{-10} = \frac{(0.30 - x)x}{0.10 - x}$ and assume $x << 0.10$ (and << 0.30)

$x = 1.4 \times 10^{-10}$

Ans: 9.85

8) Calculate the pH of a solution which is 0.010 M in NH_3.
The K_b for NH_3 is 1.8 x 10^{-5}.
$NH_3 + H_2O \rightleftharpoons OH^- + NH_4^+$
— x—>

	$[NH_3]$	$[OH^-]$	$[NH_4^+]$
initially	0.010	0	0
equilibrium	0.010 - x	x	x

substituting: $$1.8\times10^{-5} = \frac{x^2}{0.010 - x}$$ and assume $x << 0.010$

$x^2 = 1.8 \times 10^{-7}$ so: $x = 4.2 \times 10^{-4}$ (=$[OH^-]$)
pOH = 3.37

Ans: 10.63

9) Calculate the pH of a solution which is 0.010 M in NH_3 and 0.010 M in NH_4^+.
The K_b for NH_3 is 1.8 x 10^{-5}.
$NH_3 + H_2O \rightleftharpoons OH^- + NH_4^+$
— x—>

	$[NH_3]$	$[OH^-]$	$[NH_4^+]$
initially	0.010	0	0
equilibrium	0.010 - x	x	0.010 + x

substituting: $$1.8\times10^{-5} = \frac{(0.010 + x)x}{0.010 - x}$$ and assume $x << 0.010$

$x = 1.8 \times 10^{-5}$ (=$[OH^-]$)
pOH = 4.74

Ans: 9.26

10) Calculate the pH of a solution which is 0.10 M in NH_3 and 0.010 M in NH_4^+. The K_b for NH_3 is 1.8 x 10^{-5}.
$NH_3 + H_2O \rightleftharpoons OH^- + NH_4^+$
— x—>

	$[NH_3]$	$[OH^-]$	$[NH_4^+]$
initially	0.010	0	0
equilibrium	0.10 - x	x	0.010 + x

substituting: $$1.8\times10^{-5} = \frac{(0.010 + x)x}{0.10 - x}$$ and assume $x << 0.010$ (and << 0.10)

$x = 1.8 \times 10^{-4}$ (=$[OH^-]$)
pOH = 3.74

Ans: 10.26

Problems 11, 12 and 13 are dilution problems. If you do not know how to do dilution problems, do some review first. These are in preparation for subsequent questions.

11) If 100.0 mL of a solution is originally 0.10 M CH_3COOH and one adds 50.0 mL of water to this solution, what is the final concentration of the CH_3COOH?

For dilutions: $C_1V_1 = C_2V_2$

$C_1 = 0.10$ M
$V_1 = 100.0$ mL
$C_2 = ?$
$V_2 = 150$ mL (volume of dilute solutions add with little error.)

substitute:
(0.10 M)(100.0 mL) = C_2(150 mL)

Ans: 0.067 M

12) If 100.0 mL of a solution is originally 0.10 M CH_3COOH and one adds 50.0 mL of 0.10 M NaCl to this solution, what is the final concentration of the CH_3COOH?

$C_1V_1 = C_2V_2$
For CH_3COOH: (NaCl is irrelevant)
(0.10 M)(100.0 mL) = C_2(150.0 mL)

Ans: 0.067 M

13) If 100.0 mL of a solution is originally 0.10 M CH_3COOH and one adds 50.0 mL of 0.10 M $NaCH_3COO$ to this solution, what is the final concentration of the CH_3COOH? What is the final concentration of the $NaCH_3COO$? (At this point, ignore the establishment of the Brønsted-Lowry equilibrium.)

$C_1V_1 = C_2V_2$

For CH_3COOH:
(0.10 M)(100.0 mL) = C_2(150.0 mL)

$NaCH_3COO$:
(0.10 M)(50.0 mL) = C_2(150.0 mL)

For M of CH_3COOH: Ans: 0.067 M CH_3COOH

For M of $NaCH_3COO$ Ans: 0.033 M $NaCH_3COO$

(What is Na^+ in $NaCH_3COO$?) What is the important ion?

Ans: the CH_3COO^- (acetate) ion

14) What are the pHs of the solutions described in questions 11, 12 and 13? The K_a for $CH_3COOH = 1.8 \times 10^{-5}$.

For 11:

$CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$

— x—>

	$[CH_3COOH]$	$[H_3O^+]$	$[CH_3COO^-]$
initially	0.067	0	0
equilibrium	0.067 - x	x	x

substituting: $1.8\times10^{-5} = \dfrac{x^2}{0.067 - x}$ and assume $x << 0.10$ (and << 0.30)

$x^2 = 1.2 \times 10^{-6}$ $x = 1.1 \times 10^{-3}$ Ans #11: pH = 2.96

For 12: Same as for 11 since Na^+ and Cl^- are a spectator ions.

Ans #12: pH = 2.96

For 13:

$CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$

— x—>

	$[CH_3COOH]$	$[H_3O^+]$	$[CH_3COO^-]$
initially	0.067	0	0.033
equilibrium	0.067 - x	x	0.033 + x

substituting: $1.8\times10^{-5} = \dfrac{(0.033 + x)x^2}{0.067 - x}$ and assume $x << 0.0.33$ (and << 0.0.67)

$x = 3.6\times 10^{-5}$

Ans #13: pH = 4.44

15) A solution is created by mixing 100.0 mL of a solution which is 0.10 M CH_3COOH with 50.0 mL of a solution which is 0.10 M NaOH. (Confused? - Look up Arrhenius Acid/Base and neutralization.)

a) Write the Arrhenius acid-base reaction for this:

$CH_3COOH + NaOH \rightarrow NaCH_3COO + H_2O$

b) What is the concentration of the CH_3COOH if one does not take into account the Brønsted-Lowry reaction with water?

By definition C = n/V where V is in liters.

and so $n = CV$

moles of:	when:	How calculated:	ANS:
CH_3COOH	initially	(0.10 M)(0.1000 L)	0.010 mol
NaOH	initially	(0.10 M)(0.0500 L)	0.0050 mol
$NaCH_3COOH$	after Arrhenius reaction	$n_{NaCH3COO}/1= n_{NaOH}/1$*	0.0050 mol
CH_3COOH	after Arrhenius reaction	$n_{starting} - n_{used}$	0.005 mol

*limiting reactant.

thus: $C_{CH_3COOH} = \dfrac{0.0050\text{ mol}}{0.1500\text{ L}}$ Ans: 0.033 M CH_3COOH

15 continued:

c) What is the concentration of the $NaCH_3COO$ if one does not take into account the Brønsted-Lowry reaction with water?

thus: $C_{NaCH_3COO} = \dfrac{0.0050\text{ mol}}{0.1500\text{ L}}$ Ans: 0.033 M $NaCH_3COO$

d) What is the *pH* of the resultant solution described above?
The K_a for $CH_3COOH = 1.8 \times 10^{-5}$.

$pH = pK_a = -\log(1.8 \times 10^{-5}) = 4.74$ (Why?)

Ans: *pH* = 4.74

16) What is the *pH* of a solution formed by mixing 200 mL of 0.20 M NH_3 with 50 mL of water? K_b for $NH_3 = 1.8 \times 10^{-5}$.

$C_1V_1 = C_2V_2$

$C_1 = 0.20$ M
$V_1 = 200$ mL
$C_2 = ?$
$V_2 = 250$ mL
$(0.20\text{ M})(200\text{ mL}) = C_2(250\text{ mL})$ $C_2 = 0.16$ M

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$$1.8\times10^{-5} = \frac{x^2}{0.16 - x}$$

$x^2 = 2.9 \times 10^{-6}$ $x = 1.6 \times 10^{-3}$ (= $[OH^-]$) Ans: pH = 11.23

17) What is the pH of a solution formed by mixing 200 mL of 0.20 M NH_3 with 50 mL of 0.50 M NH_4Cl? The K_b for NH_3 = 1.8 x 10^{-5}.

$C_1V_1 = C_2V_2$ for both [NH_3] and NH_4^+

[NH_3]:	[NH_4^+]:
C_1 = 0.20 M	C_1 = 0.50 M
V_1 = 200 mL	V_1 = 50 mL
C_2 = ?	C_2 = ?
V_2 = 250 mL	V_2 = 250 mL
C_2 = 0.16 M	C_2 = 0.10 M

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} \qquad 1.8\times10^{-5} = \frac{(0.10+x)x}{0.16-x}$$

$x = 2.9 \times 10^{-5}$ (= [OH^-]) pOH = 4.54 Ans: *pH* = 9.46

18) What is the *pH* of a solution formed by mixing 200 mL of 0.20 M NH_3 with 50 mL of 0.50 M HCl? (Note: you must first answer questions similar to question ? above. What is the concentration of NH_3 and NH_4Cl?) The K_b for NH_3 = 1.8 x 10^{-5}.

Final volume = 250 mL

n_{NH3} originally = (0.20 M)(200 mL) = 40 mmol

n_{HCl} originally = (0.50 M)(50 mL) = 25 mmol (limiting reactant)

n_{NH4Cl} produced = n_{HCl} originally = 25 mmol

n_{NH3} left = 40 mmol – 25 mmol = 15 mmol

resultant [NH_3] = 15 mmol/250 mL = 0.060

resultant [NH_4^+] = 25 mmol/250 mL = 0.10

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} \qquad 1.8\times10^{-5} = \frac{(0.10+x)x}{0.060-x}$$

$x = 1.1 \times 10^{-5}$ (=[OH^-]) pOH = 4.97 Ans: pH = 9.03

19) What is the pH of a solution formed by mixing 200 mL of 0.20 M NH_3 with 50 mL of 0.50 M NH_4Cl and 50 mL of 0.50 M HCl? (Hint: first calculate the dilutions for each of these species, then do the Arrhenius acid-base reaction.) The K_b for NH_3 = 1.8 x 10^{-5}.

Final Volume = 300 mL

n_{NH3} originally = (0.20 M)(200 mL) = 40 mmol

n_{HCl} originally = (0.50 M)(50 mL) = 25 mmol (limiting reactant)

n_{NH4Cl} originally = (0.50 M)(50 mL) = 25 mmol

n_{NH4Cl} produced = n_{HCl} originally = 25 mmol

n_{NH4Cl} total = 25 mmol + 25 mmol = 50 mmol

n_{NH3} left = 40 mmol – 25 mmol = 15 mmol

resultant [NH_3] = 15 mmol/300 mL = 0.050

resultant [NH_4^+] = 50 mmol/300 mL = 0.167

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} \qquad 1.8\times10^{-5} = \frac{(0.167+x)x}{0.050-x}$$

$x = 5.4 \times 10^{-6}$ (=[OH^-]) pOH = 5.27 Ans: pH = 8.73

Weak Acids, Weak Bases and Buffers

Use the following K_as to answer questions 20 through 23:

For the reaction:	The equilibrium constant is:
$H_3PO_4 + H_2O \rightleftharpoons H_2PO_4^- + H_3O^+$	$K_{a1}(H_3PO_4) = 7.5 \times 10^{-3}$
$H_2PO_4^- + H_2O \rightleftharpoons HPO_4^{2-} + H_3O^+$	$K_{a2}(H_2PO_4^-) = 6.2 \times 10^{-8}$
$HPO_4^{2-} + H_2O \rightleftharpoons PO_4^{3-} + H_3O^+$	$K_{a3}(HPO_4^{2-}) = 3.6 \times 10^{-13}$

20) What is the pH of a solution which is 0.10 M in H_3PO_4?

$$K_{a1} = \frac{[H_3PO_4][H_3O^+]}{[H_2PO_4^-]}$$

$$7.5\times10^{-3} = \frac{x^2}{0.10 - x}$$

etc.

Ans: *pH* = 1.56 (1.62 if x << 0.10 not assumed)

21) What is the pH of a solution which is 0.10 M in H_3PO_4 and 0.10 M in NaH_2PO_4?

(What is Na^+? It is spectator ion)

pH = pK_{a1}

Ans: *pH* = 2.12

22) What is the pH of a solution which is 0.10 M in NaH_2PO_4 and 0.10 M in Na_2HPO_4?

pH = pK_{a2}

Ans: *pH* = 7.21

23) What is the *pH* of a solution which is 0.10 M in Na_2HPO_4 and 0.10 M in Na_3PO_4?

pH = pK_{a3}

Ans: *pH* = 12.44